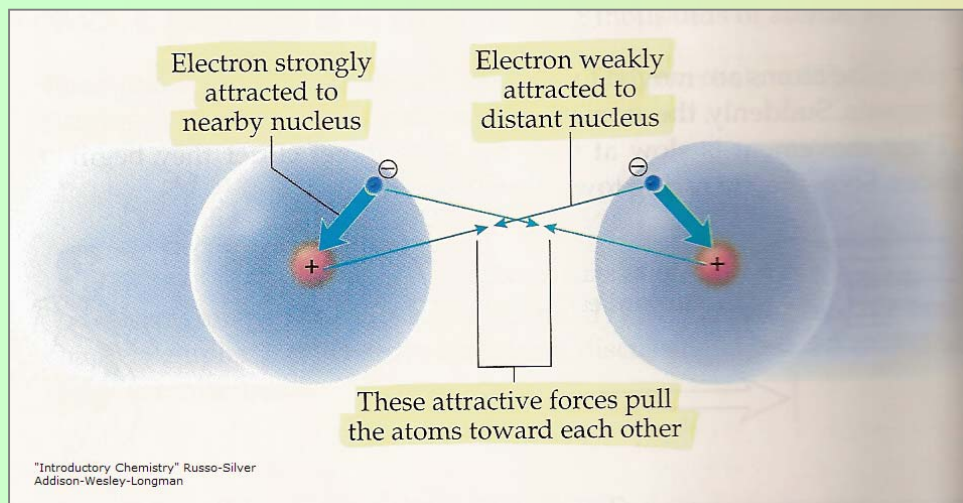


# Covalent Bonding



## The Formation of a Covalent Bond

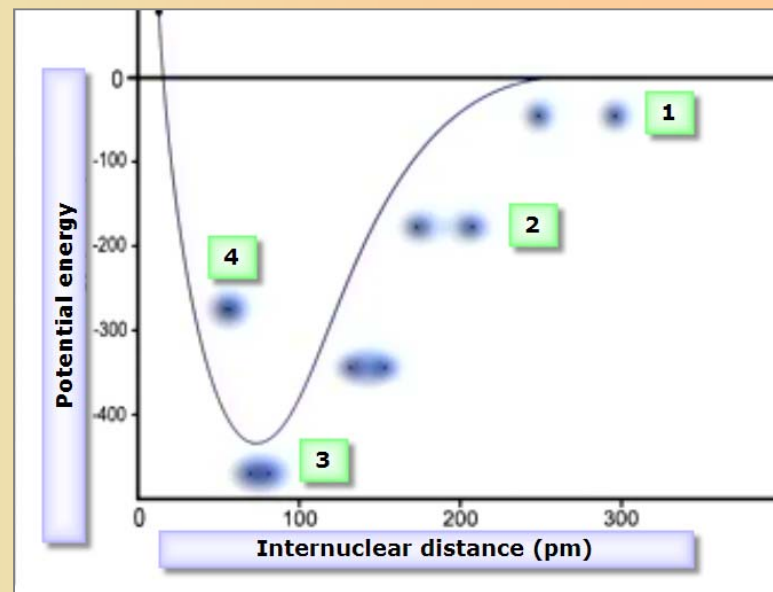
When two isolated H atoms approach each other they tend to form a  $H_2$  molecule.

At "1" the atoms are too far apart to attract each other

• At "2" each nucleus attracts the other atom's electron.

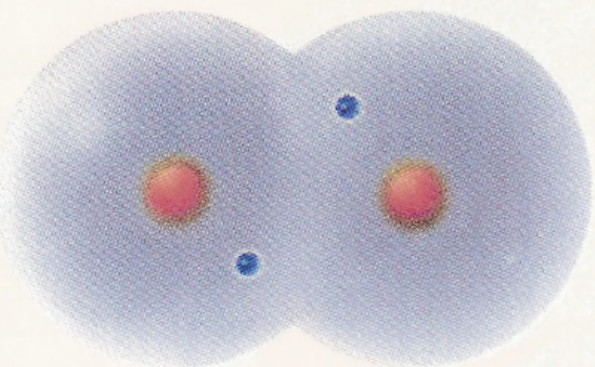
• At "3" the combination of nucleus-electron attraction and electron-electron and nucleus-nucleus repulsions gives the minimum energy of the system

• At "4" repulsions increase the system's energy and force the atoms to point "3" again

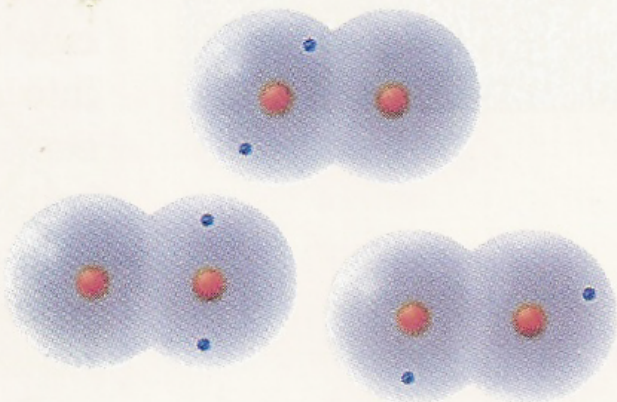


# Covalent Bonding

The electrons spend most of their time between nuclei . . .



although they spend some time in other places



"Introductory Chemistry" Steve Russo, Mike Silver  
Addison-Wesley-Longman

In a covalent bonding the electrons are shared between two nuclei (atoms).

We can use the Lewis notation to describe the bonding in this way:

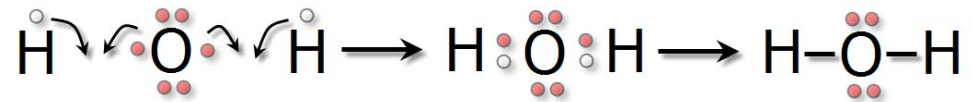


A shared pair of electrons can be represented by a line:



# Covalent Bonding

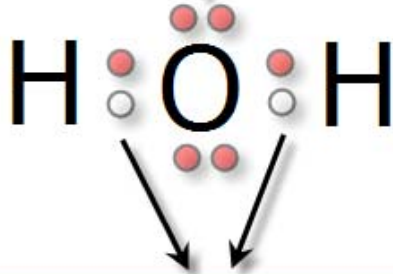
Oxygen and hydrogen can combine together, share electrons (covalent bonds) and form a molecule:



From this example, we can draw the following conclusions:

- Only outer level electrons take part in covalent bonds
- Only unpaired electrons (from this outer level) are shared and form covalent bonds
- In a molecule, we have two types of electrons: shared electrons and lone electrons (not shared)

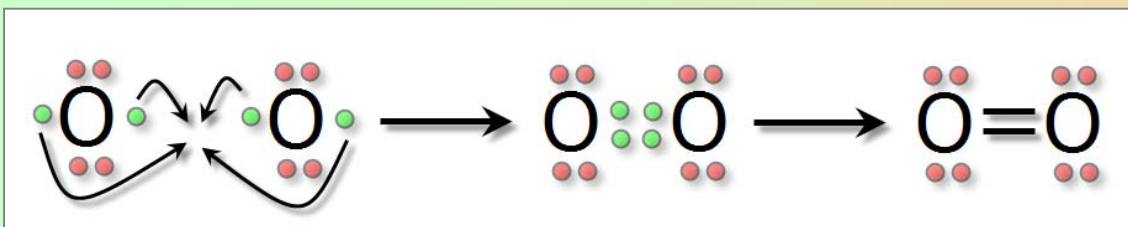
lone pair:  
they don't take part in covalent bonds



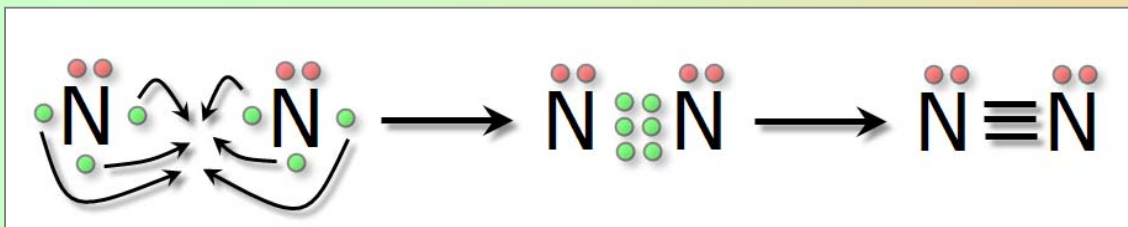
bonding pair:  
they take part in covalent bonds

# Covalent Bonding

A **double bond** consists of two bonding pairs, four electrons shared between two atoms.

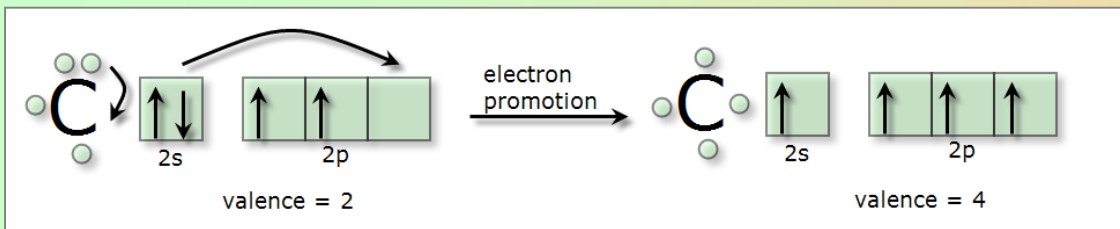


A **triple bond** consists of three bonding pairs, six electrons shared between two atoms.

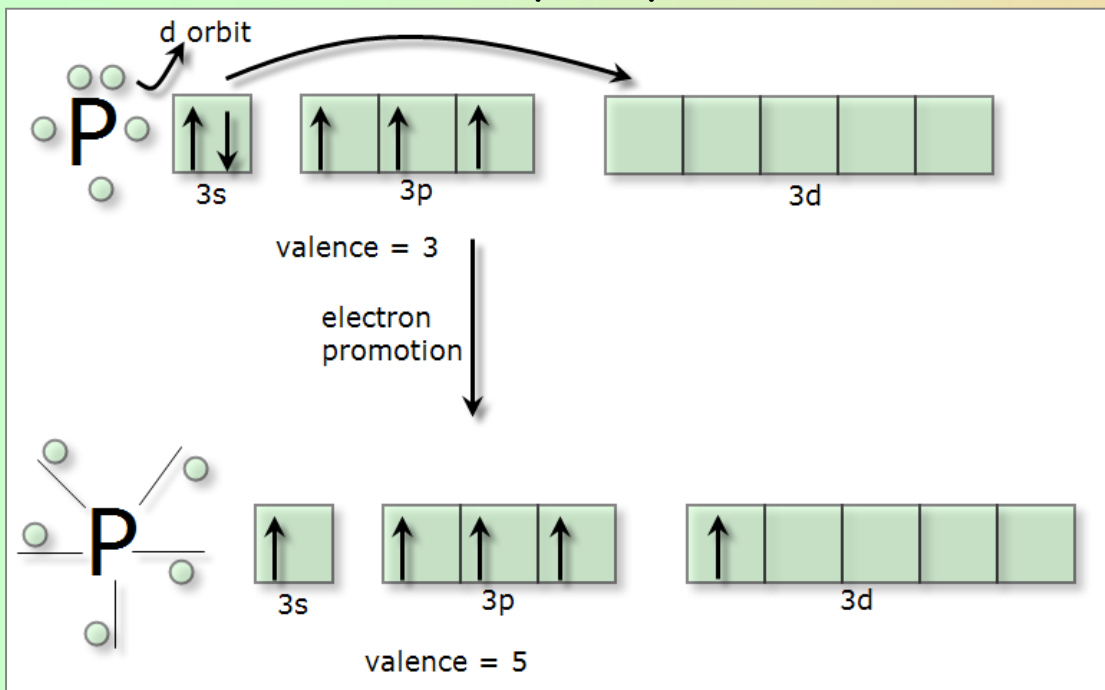


# Covalent Bonding

## Covalent valences of carbon: 2, 4



## Covalent valences of phosphorus: 3, 5



## Covalent valence

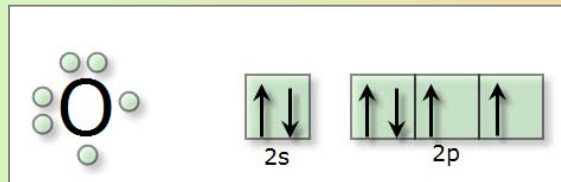
We will define "covalent valence" as the number of unpaired electrons an atom can have.

In order to have more unpaired electrons the outer electrons of an atom can have a promotion: they can jump from one orbit to another always at the same level (they can not jump from  $n=2$  to  $n=3$ )

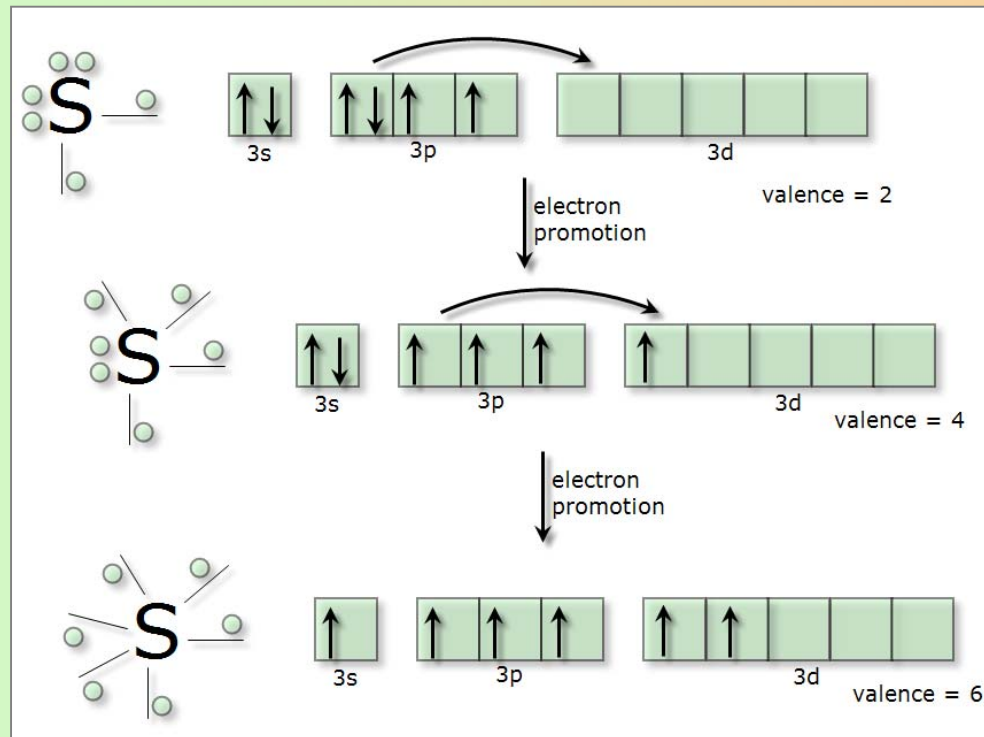
# Covalent Bonding

Diference between oxygen (valence 2) and sulphur (2, 4, 6)

Oxygen can't promote more electrons, because there is no "2d" orbits:



But sulphur (which belongs to the same group) can promote its electrons:



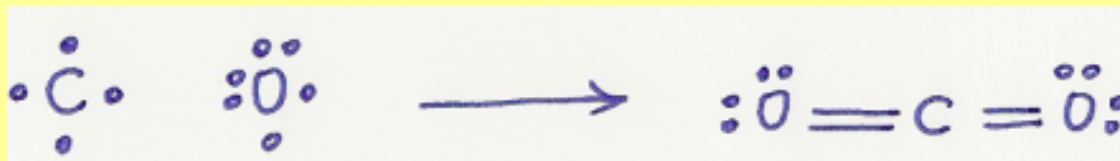
# Covalent Bonding

**EXERCISE:** Draw the following molecules, using the Lewis notation

NH<sub>3</sub>



CO<sub>2</sub>



CH<sub>4</sub>


H<sub>2</sub>S

# Covalent Bonding

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<b>CO</b>	
<b>BF<sub>3</sub></b>	
<b>HCl</b>	

# Covalent Bonding

SO <sub>2</sub>	 <p>The diagram shows the Lewis structure of sulfur dioxide (SO<sub>2</sub>). On the left, a sulfur atom (S) is shown with four single bonds (two to the left and two to the right) and one lone pair on the left. To its right is an oxygen atom (O) with two single bonds (one above and one below) and two lone pairs. An arrow points to the right, where the final Lewis structure is shown: sulfur is double-bonded to two oxygen atoms. Each oxygen atom has two lone pairs. The sulfur atom has no lone pairs.</p>
SO <sub>3</sub>	
PCl <sub>3</sub>	
PCl <sub>5</sub>	

# Covalent Bonding

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$\text{SF}_6$	
$\text{HCN}$	
$\text{P}_2\text{O}_5$	
$\text{Cl}_2\text{O}_5$	